

OCR A Chemistry A-Level

Module 3 - Periodic Table and Energy

Enthalpy Changes

Notes and Example Calculations

Answers given at the end of the booklet





Calorimetry

Enthalpy changes cannot be measured directly but can be measured by monitoring the temperature of a chemical system. The units for enthalpy change is kJmol^{-1} . Calorimetry uses a mathematical relationship to work out the enthalpy change from experimental data. The formula used is:

$$q = mc\Delta T$$

(q = energy change) (m = mass) (c = specific heat capacity) (ΔT = temperature change)

Example 1:

How much energy is needed to change the temperature of 60.0 g of water by 15.0 °C.

Step 1: Input the values into the formulae. (Specific heat capacity = $4.18 \text{ Jg}^{-1}\text{K}^{-1}$)

$$\begin{aligned}\Rightarrow q &= 60 \times 4.18 \times 15 \\ &= \mathbf{3762 \text{ J}} \\ &= \mathbf{3.76 \text{ kJ}}\end{aligned}$$

Example 2:

0.25 g of ethene was used to heat 200 cm^3 of water from 20 °C to 42.2 °C. What was the enthalpy change of this reaction?

[Remember 1 cm^3 of water = 1g]

Step 1: Calculate the change in temperature.

$$\Rightarrow 42.2 - 20 = 22.2$$

Step 2: Work out the energy released in this reaction

$$\begin{aligned}\Rightarrow q &= 200 \times 4.18 \times 22.2 \\ &= 1.856 \times 10^4 \text{ J} \\ &= 18.56 \text{ kJ}\end{aligned}$$

Step 3: Calculate the number of moles of ethane (C_2H_6)

$$\text{Mr of ethene} = 12 + 12 + 6 = 30$$

$$\Rightarrow 0.25 / 30 = 0.00833$$





Step 3: Divide the energy released by the moles of ethane used to calculate the enthalpy change per mole.

$$\Rightarrow 18.56 / 0.00833 = 2236 \text{ kJ/mol}$$

As this is an **exothermic** reaction, enthalpy change is **negative** therefore the final answer is

-2236 kJmol⁻¹

Worked exam style question

Question 1

A student carries out an experiment to determine the enthalpy change of combustion of glucose.

In the experiment, 0.831 g of glucose is burned. The energy released is used to heat 100 cm³ of water from 23.7 °C to 41.0 °C.

(i) Calculate the energy released, in kJ, during combustion of 0.831 g glucose.

The specific heat capacity of water = 4.18 J g⁻¹ K⁻¹.

Density of water = 1.00 g cm⁻³.

(ii) Calculate the amount, in moles, of glucose that is burned.

(iii) Calculate the enthalpy change of combustion of glucose.
Give your answer to **three** significant figures.

Step 1: Use $q = mc\Delta t$ to calculate the energy released.

$$\begin{aligned}\Rightarrow q &= 100 \times 17.3 \times 4.18 \\ &= \underline{\underline{7.231 \text{ kJ}}}\end{aligned}$$

Step 2: Use $\text{moles} = \text{mass} / M_r$ to calculate the moles of glucose.

M_r of glucose = 180

$$\Rightarrow 0.831 / 180 = \underline{\underline{0.00462 \text{ mol}}}$$

Step 3: Use the previous answers to find enthalpy change per mole.

$$\Rightarrow 7.2314 / 0.00462 = \underline{\underline{1570 \text{ kJmol}^{-1}}}$$





Try these questions...

1. A student carried out an experiment to determine the enthalpy change of combustion of pentan-1-ol, $\text{CH}_3(\text{CH}_2)_4\text{OH}$.

In the experiment, 1.76 g of pentan-1-ol was burnt. The energy was used to heat 250 cm^3 of water from $24.0 \text{ }^\circ\text{C}$ to $78.0 \text{ }^\circ\text{C}$.

- (i) Calculate the energy released, in kJ, during combustion of 1.76 g pentan-1-ol.

The specific heat capacity of water = $4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

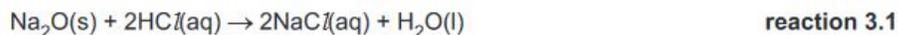
Density of water = 1.00 g cm^{-3} .

- ii) Calculate the amount, in moles, of pentan-1-ol that was burnt.
(iii) Calculate the enthalpy change of combustion of pentan-1-ol.

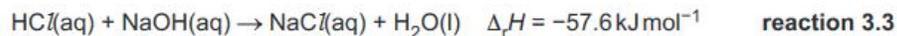
[6 marks]

Q2.

A student plans to determine the enthalpy change of **reaction 3.1** shown below.



This enthalpy change can be determined indirectly using Hess' Law from the enthalpy changes of **reaction 3.2** and **reaction 3.3** shown below.



The student will determine the enthalpy change of **reaction 3.2** as outlined below.

- Weigh a bottle containing $\text{Na}_2\text{O}(\text{s})$ and weigh a polystyrene cup.
- Add about 25 cm^3 of water to the polystyrene cup and measure its temperature.
- Add the $\text{Na}_2\text{O}(\text{s})$, stir the mixture, and measure the maximum temperature reached.
- Weigh the empty bottle and weigh the polystyrene cup with the final solution.

Mass readings

Mass of bottle + $\text{Na}_2\text{O}(\text{s})$ = 16.58 g
Mass of empty bottle = 15.34 g

Mass of empty polystyrene cup = 21.58 g
Mass of polystyrene cup + final solution = 47.33 g

Temperature readings

Initial temperature of water = $20.5 \text{ }^\circ\text{C}$
Maximum temperature of final solution = $55.5 \text{ }^\circ\text{C}$

)* Calculate the enthalpy change of **reaction 3.2** and the enthalpy change of **reaction 3.1**.

Show all your working.





[6 marks]

Bond Enthalpies

Exam definition: the energy required to break 1 mole of bonds

The enthalpy change for a reaction can be calculated from average bond enthalpy data using the following expression:

$$\Delta H = \Sigma (\text{bond enthalpies of reactants}) - \Sigma (\text{bond enthalpies of products})$$

bond	$\Delta H_b/\text{kJmol}^{-1}$	bond	$\Delta H_b/\text{kJmol}^{-1}$
C-H	+413	H-F	+565
C-Br	+280	C-F	+425
H-Br	+366	F-F	+158
Br-Br	+193	C=O	+805
C-C	+347	O-H	+464
C=C	+611	O=O	+498
H-H	+435		

Example 1:

The ΔH for the following reaction: $\text{CH}_4(\text{g}) + \text{Br}_2(\text{g}) \rightarrow \text{CH}_3\text{Br}(\text{g}) + \text{HBr}(\text{g})$

Step 1: For the reactants write out the type and number of each of the bonds.

\Rightarrow C - H (x4)
Br - Br (x1)

Step 2: Calculate the sum of the bond enthalpies of reactants.

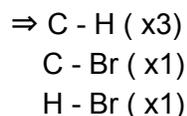
$\Rightarrow 4 \times 413$ (C-H) = 1652
 1×193 = 193

$\Rightarrow 1652 + 193 = 1845$

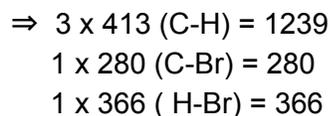




Step 3: For the products write out the type and number of each of the bonds.



Step 4: Calculate the sum of the bond enthalpies of products.



$$\Rightarrow 1239 + 280 + 366 = 1885$$

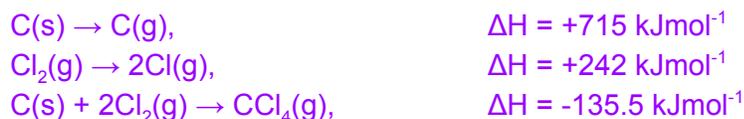
Step 5: Calculate the enthalpy change of the reaction

$$[\Sigma (\text{bond enthalpies of reactants}) - \Sigma (\text{bond enthalpies of products})]$$

$$1845 - 1885 = \underline{\underline{-40 \text{ kJ mol}^{-1}}}$$

Example 2:

Given the following information:



Calculate the average bond dissociation enthalpy of a C-Cl bond.

Step 1: Find the total bond enthalpies for the reactants.

\Rightarrow To do this you have to compare the first two equations with the reactants of the third equation.

[There is one carbon atom in the first equation and in the third so the bond enthalpy can be taken from the first equation. However the third equation has 2 chlorine molecules whereas the second only has one. Therefore multiply the enthalpy change of the second reaction by 2 to get the bond enthalpy for 2 chlorine molecules.]

$$\Rightarrow 715 + (242 \times 2) = 1199$$

Step 2: Rearrange the equation so that the sum of bond enthalpies of products is the subject,



to work out the bond enthalpy of a C-Cl bond.

$$\Sigma (\text{bond enthalpies of reactants}) - \Delta H = \Sigma (\text{bond enthalpies of products})$$

$$\Rightarrow 1199 - (-135.5) = 1334.5$$

$$[1334.5 = 4 \times \text{C-Cl}]$$

Step 3: Calculate 1 C-Cl bond.

$$1334.5 / 4 = \underline{\underline{+333.6 \text{ J}}}$$

[Remember bond enthalpies are always **positive** because bond breaking is an **endothermic** process]

Worked Exam Style Question

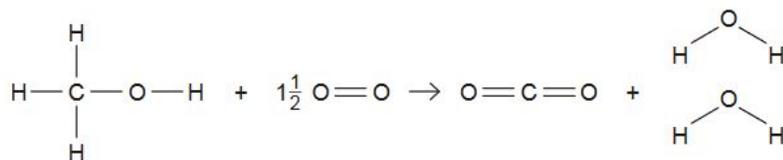
Question 1

Energy changes during reactions can be considered using several different enthalpy changes. These include average bond enthalpies and enthalpy changes of combustion.

Table below shows the values of some average bond enthalpies.

bond	average bond enthalpy / kJ mol^{-1}
C — H	+ 410
O — H	+ 465
O == O	+ 500
C == O	+ 805
C — O	+ 336

- (iii) The equation below shows the combustion of methanol, CH_3OH , in the gaseous state.



Use the average bond enthalpies from the table above to calculate the enthalpy change of combustion of gaseous methanol, ΔH_c .

$$\Delta H = \quad \quad \quad \text{kJ mol}^{-1}$$

[3]



Step 1: Find the total for the bond enthalpies of the reactants.

$$\Rightarrow 3 \times 410 \text{ (C-H)} = 1230$$

$$1 \times 336 \text{ (C-O)} = 336$$

$$1 \times 465 \text{ (O-H)} = 465$$

$$1.5 \times 500 \text{ (O=O)} = 750$$

$$1230 + 336 + 465 + 750 = 2781$$

Step 2: Find the total for the bond enthalpies of the products.

$$\Rightarrow 2 \times 805 \text{ (C=O)} = 1610$$

$$4 \times 465 \text{ (O-H)} = 1860$$

$$1610 + 1860 = 3470$$

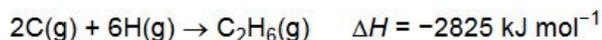
Step 3: Calculate the enthalpy change of combustion.

$$\Rightarrow 2781 - 3470 = \underline{\underline{-689 \text{ kJmol}^{-1}}}$$

Try these questions...

3.

The equations for the combination of gaseous atoms of carbon and hydrogen to form methane, CH_4 , and ethane, C_2H_6 , are shown below.



Use these data to calculate:

(i) the bond enthalpy of a C-H bond,

bond enthalpy = kJ mol^{-1}

[1]

(ii) the bond enthalpy of a C-C bond.

bond enthalpy = kJ mol^{-1}

[2]

[Total 3 marks]

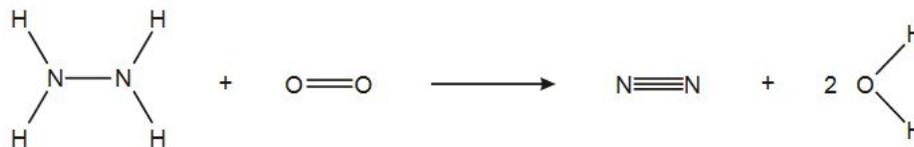




4.

This question is about hydrazine, N_2H_4 , and ammonia, NH_3 . These are both compounds of nitrogen and hydrogen.

- (a) Hydrazine can be oxidised and used as a rocket fuel. The equation for one possible reaction taking place is shown below.



Some average bond enthalpies are given below.

bond	bond enthalpy / kJ mol^{-1}
N-N	+163
$\text{N}\equiv\text{N}$	+945
N-H	+390
O=O	+497
O-H	+463

- (i) Use these data to calculate the enthalpy change for the reaction of hydrazine with oxygen, as shown.

answer kJ mol^{-1}

[4]

- (ii) Calculate the enthalpy change for one gram of hydrazine in this reaction.

answer kJ

[1]



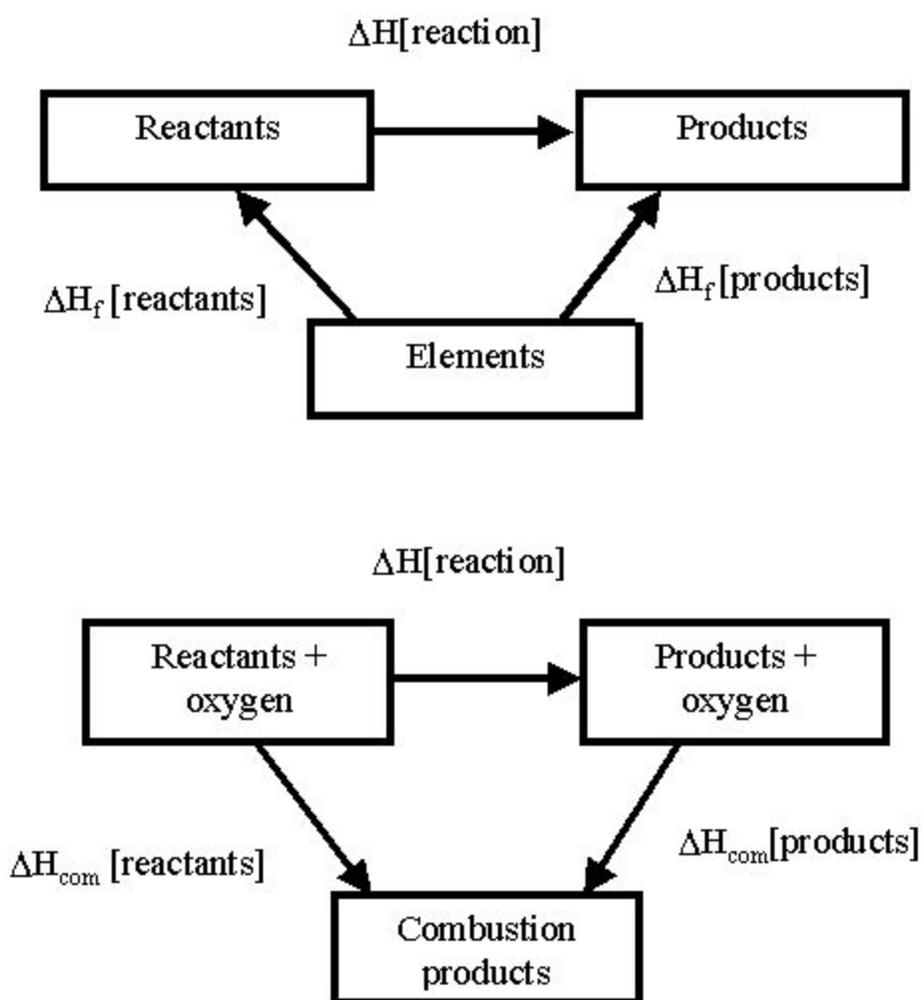


Hess' Law and Enthalpy Cycles

Enthalpy change of formation is the enthalpy change when 1 mole of a substance is formed from its constituent elements.

Enthalpy change of combustion is the enthalpy change when 1 mole of substance is completely burned in excess oxygen.

Some enthalpy changes are impossible to measure directly and for these reactions Hess' cycles can be used to indirectly calculate the enthalpy change.





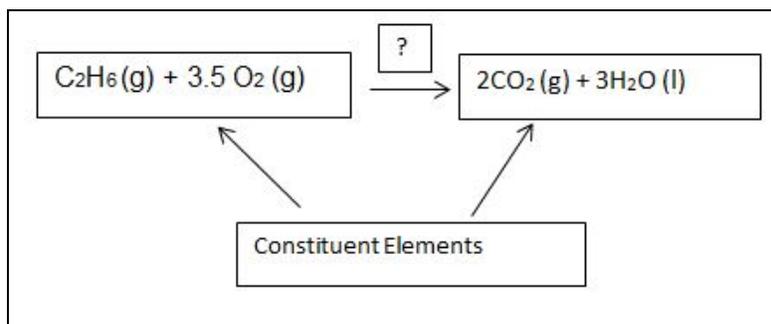
Example 1:

Substance	H ₂ O(l)	CO ₂ (g)	Ethane C ₂ H ₆ (g)	Ethene C ₂ H ₄ (g)
$\Delta H_f / \text{kJmol}^{-1}$	-285.5	-393	-83.6	+52.0

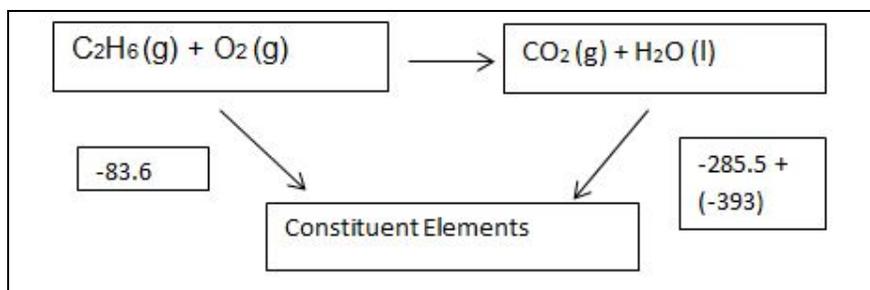
Calculate the enthalpy of combustion of ethane. The equation is below.



Step 1: Look at the information given in the question, as the information is enthalpy change of formation you know which type of Hess cycle to draw.

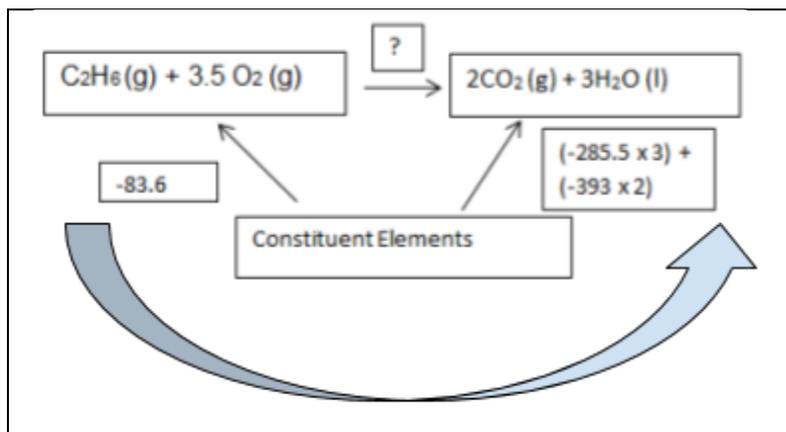


Step 2: Input values for the corresponding substances into the Hess' cycle.





Step 3: Calculate the enthalpy change of combustion by following arrows of the cycle.



$$\Rightarrow -(-83.6) + ((-285.5 \times 2) + (-393 \times 3)) = \underline{1558.9 \text{ kJmol}^{-1}}$$

Example 2:

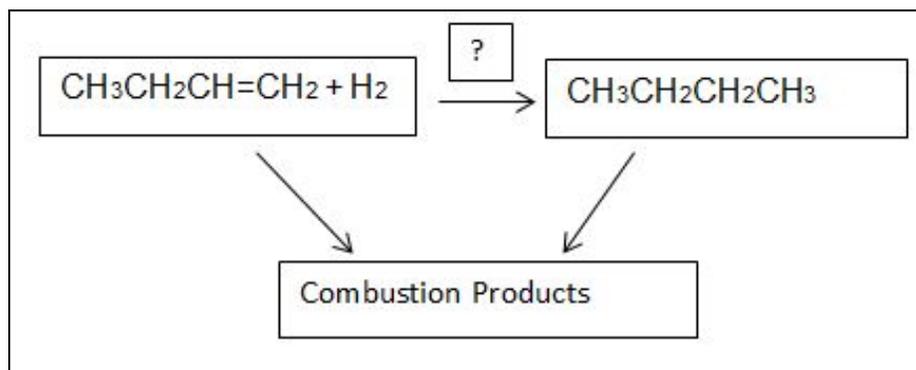
Given the following data:

Substance	CH ₃ CH ₂ CH ₂ CH ₃	CH ₃ CH ₂ CH=CH ₂	H ₂
$\Delta H_c / \text{kJmol}^{-1}$	-2877	-2717	-286

Calculate ΔH for the following reaction:

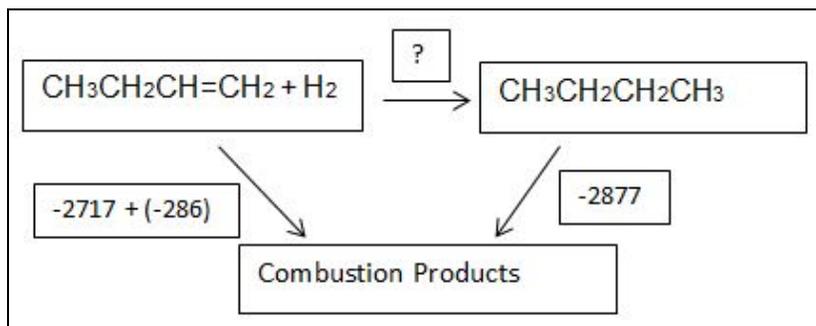


Step 1: Look at the information given in the question, as the information is enthalpy change of combustion you know which type of Hess cycle to draw.

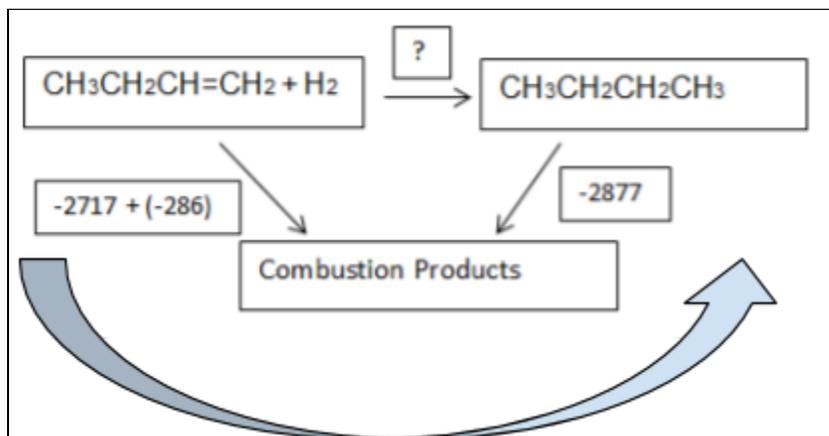




Step 2: Input values for the corresponding substances into the Hess' cycle.



Step 3: Calculate the enthalpy change of combustion by following arrows of the cycle.



$$\Rightarrow (-2717 + (-286)) - (-2877) = \underline{-126 \text{ kJmol}^{-1}}$$

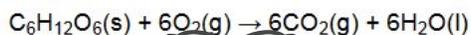
Worked Exam Style Questions

Question 1

The standard enthalpy change of combustion of glucose can also be determined indirectly.

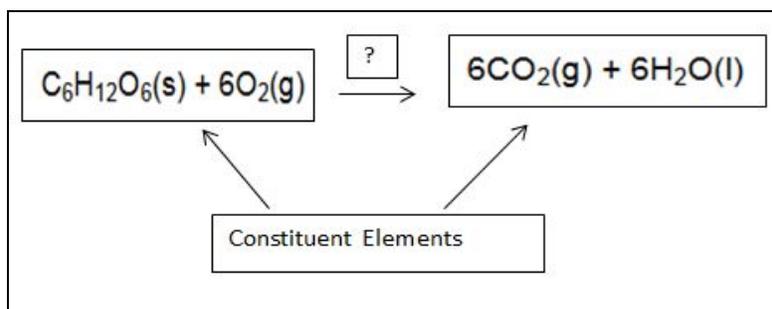
Calculate the standard enthalpy change of combustion of glucose using the standard enthalpy changes of formation below.

substance	$H_f^\ominus / \text{kJ mol}^{-1}$
$\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$	-1250
$\text{CO}_2(\text{g})$	-394
$\text{H}_2\text{O}(\text{l})$	-286

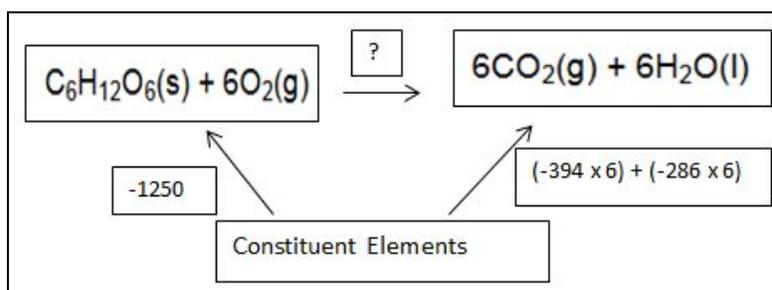




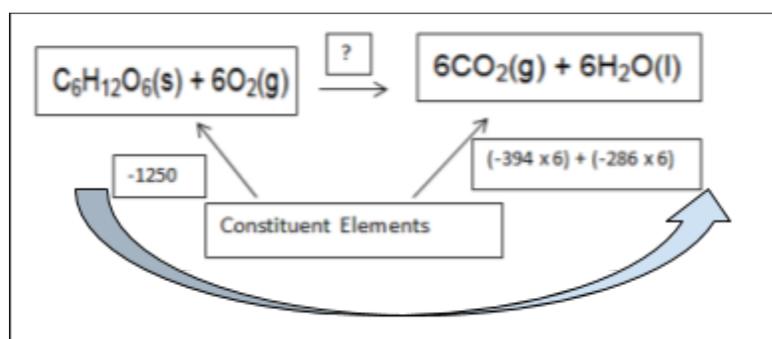
Step 1: Look at the information given in the question, as the information is enthalpy change of formation you know which type of Hess cycle to draw.



Step 2: Input values for the corresponding substances into the Hess' cycle.



Step 3: Calculate the enthalpy change of combustion by following arrows of the cycle.



$$\Rightarrow -(-1250) + ((-394 \times 6) + (-286 \times 6)) = \underline{\underline{-2830 \text{ kJmol}^{-1}}}$$

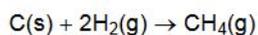


Question 2

Enthalpy changes can be calculated using enthalpy changes of combustion. The table below shows some values for standard enthalpy changes of combustion.

substance	$\Delta H_c^\ominus / \text{kJ mol}^{-1}$
C(s)	-394
H ₂ (g)	-286
CH ₄ (g)	-890

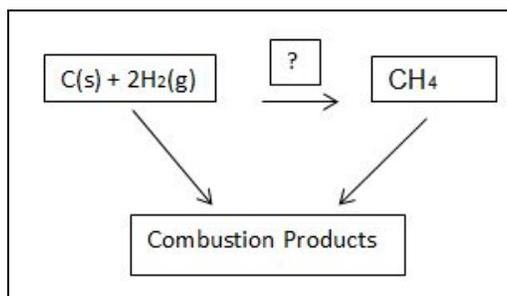
Use these values to calculate the standard enthalpy change of the reaction below.



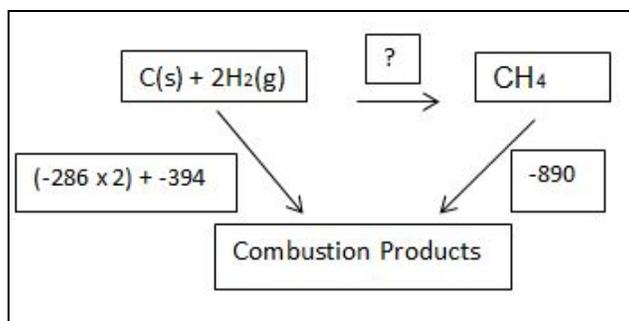
standard enthalpy change = kJ mol⁻¹

[Total 3 marks]

Step 1: Look at the information given in the question, as the information is enthalpy change of formation you know which type of Hess cycle to draw.

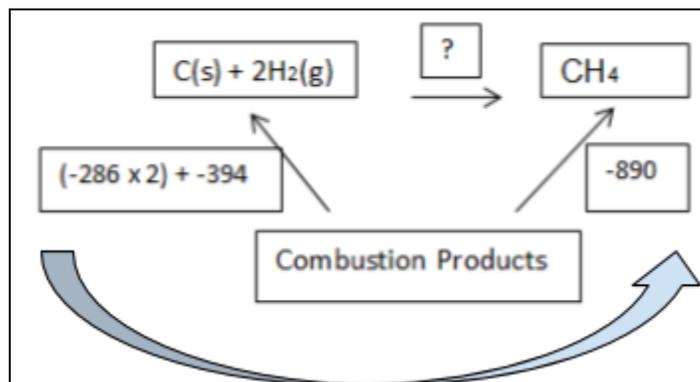


Step 2: Input values for the corresponding substances into the Hess' cycle.





Step 3: Calculate the enthalpy change of combustion by following arrows of the cycle.

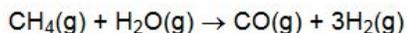


$$\Rightarrow ((-286 \times 2) + (-394)) - (-890) = \underline{-76 \text{ kJmol}^{-1}}$$

Try these questions...

5.

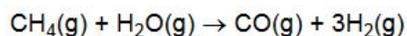
Methane reacts with steam to produce carbon monoxide and hydrogen. The equation for this process is given below.



The table below shows the standard enthalpy changes of formation for CH_4 , H_2O and CO .

compound	$\Delta H_f^\circ / \text{kJ mol}^{-1}$
CH_4	-75
H_2O	-242
CO	-110

- (iv) Using the ΔH_f° values in the table above, calculate the enthalpy change for the reaction of methane with steam.



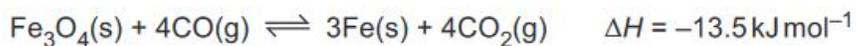
3 marks





6.

Another equilibrium involved in the extraction of iron from Fe_3O_4 is shown below.



Enthalpy changes of formation, $\Delta_f H$, for $\text{Fe}_3\text{O}_4(\text{s})$ and $\text{CO}_2(\text{g})$ are shown in the table.

Compound	$\Delta_f H / \text{kJ mol}^{-1}$
$\text{Fe}_3\text{O}_4(\text{s})$	-1118.5
$\text{CO}_2(\text{g})$	-393.5

Calculate the enthalpy change of formation, $\Delta_f H$, for $\text{CO}(\text{g})$.

3 marks





Answers

Q1.

(i) 56.430 (kJ) ✓

(ii) $M_r [\text{CH}_3(\text{CH}_2)_4\text{OH}] = 88.0$ ✓

$n = 0.0200$ mol ✓

ALLOW 88

ALLOW 0.02 **OR** ecf from wrong M_r

ALLOW full marks for 0.02 with no working out

2

(iii) (-2821.5) ✓

$= (-2820)$ (3 SF) ✓

correct minus sign ✓

ALLOW correct substitution into formula (b)(i) ÷ (b)(ii) e.g.

$56.4 \div 0.02$ this is essentially a mark for the working

ALLOW ecf from i.e. answer from (b)(i) ÷ (b)(ii)

The minus mark is stand alone and is independent of the numerical answer

3

[6]

Q2.

Please refer to the marking instructions on page 5 of this mark scheme for guidance on how to mark this question.

Level 3 (5–6 marks)

A comprehensive conclusion, using all quantitative data, to calculate the energy change and ΔH values for reactions 3.1 and 3.2

AND linking ΔH data using Hess' Law

There is a well-developed line of reasoning which is clear and logically structured. The working throughout is clearly shown. All values calculated with reasonable numbers of SF and correct signs mostly shown, allowing for ECF.

6

Indicative scientific points may include:

1. Masses and ΔT from raw results

- $m(\text{Na}_2\text{O}) = 1.24$ (g)
- $m(\text{solution}) = 25.75$ (g)
- $\Delta T = 35.0$ ($^{\circ}\text{C}$)

Energy change from $mc\Delta T$

- energy released in J **OR** kJ
 $= 25.75 \times 4.18 \times 35.0$
 $= 3767$ (J) **OR** 3.767 (kJ)
(3.767225 unrounded)





Level 2 (3–4 marks)

Attempts to describe all three scientific points but explanations may be incomplete.

OR Explains two scientific points thoroughly with few omissions.

There is a line of reasoning with some logical structure. There may be minor errors in energy change and errors in the calculations of ΔH for reaction 3.1 or reaction 3.2.

Level 1 (1–2 marks)

Processes raw mass and temperature data and obtains a calculated value for the energy change using $mc\Delta T$

OR attempts to obtain values for two scientific points but explanations may be incomplete

There is an attempt at a logical structure with a line of reasoning to obtain a value for energy change. There may be minor errors in calculation of energy change.

0 marks – No response or no response worthy of credit.

2. $\Delta_r H$ for reaction 3.2

- $n(\text{Na}_2\text{O}) = \frac{1.24}{62.0} = 0.0200 \text{ (mol)}$
- $\Delta_r H \text{ value} = \frac{3767}{0.0200} = -188 \text{ (kJ mol}^{-1}\text{)}$
(–188.36125 unrounded)

3. $\Delta_r H$ for reaction 3.1

- ΔH value for **reaction 3.1** clearly linked to ΔH for **reaction 3.2** and **reaction 3.3** in energy cycle or an expression:
 $\Delta H(3.1) = \Delta H(3.2) + 2\Delta H(3.3)$
- $\Delta H(3.1) = -188 + (2 \times -57.6)$
 $= -188 - 115.2 = -303.2 \text{ (kJ mol}^{-1}\text{)}$
(–303.56125 unrounded)

Note

Throughout, **ALLOW ECF** from previous value
ALLOW omission of trailing zeroes

Q3.

- (i) $1652/4 = 413 \text{ (kJ mol}^{-1}\text{)} \checkmark$ 1
- (ii) $(\text{C} \begin{smallmatrix} \text{O} \\ \text{O} \end{smallmatrix} \text{C}) + 6(\text{C} \begin{smallmatrix} \text{O} \\ \text{O} \end{smallmatrix} \text{H}) = 2825 \checkmark$
 $(\text{C} \begin{smallmatrix} \text{O} \\ \text{O} \end{smallmatrix} \text{C}) = 2825 - 6(413) = 347 \text{ (kJ mol}^{-1}\text{)} \checkmark$ 2

[3]

Q4.

- (a) (i) bonds broken
 $(\text{N} - \text{N}) + (\text{O} = \text{O}) + (\text{N} - \text{H}) = 163 + 497 + 4(390) = 2220 \text{ (kJ mol}^{-1}\text{)} (1)$
bonds made
 $(\text{N} \equiv \text{N}) + 4(\text{O} - \text{H}) = 945 + 4(463) = 2797 \text{ (kJ mol}^{-1}\text{)} (1)$
broken ΔH is +ve and made ΔH is –ve (1)
enthalpy of reaction $= 577 \text{ (kJ mol}^{-1}\text{)} (1)$ 4
- (ii) $\frac{577}{32} = 18.0 \text{ (kJ)} (1)$ 1

Q5.

cycle drawn/sum of $\Delta H(\text{products}) - \Delta H(\text{reactants}) (1)$

$-75 - 242 + x = -110 (1)$

$\Delta H = (+)207 \text{ kJ mol}^{-1} (1)$

3





Q6.

Answer	Marks	Guidance
FIRST, CHECK THE ANSWER ON ANSWER LINE IF answer = -110.5, Award 3 marks.	3	For answer, ALLOW -111 (kJ mol⁻¹)
Correct expression $-13.5 = (4 \times -393.5) - (-1118.5 + 4 \times \Delta_f H(\text{CO}))$ ✓		
Correct subtraction using ΔH and $\Delta_f H(\text{Fe}_3\text{O}_4)$ $4 \times \Delta_f H(\text{CO}) = (4 \times -393.5) - (-1118.5) + 13.5$ $= -442(.0)$ (kJ mol ⁻¹) ✓		NOTE: IF any values are omitted, DO NOT AWARD any marks. e.g. -393.5 OR -13.5 may be missing
Calculation of $\Delta_f H(\text{CO})$ formation $\Delta_f H(\text{CO}) = -\frac{442}{4} = -110.5 \text{ (kJ mol}^{-1}\text{)} \checkmark$		Common errors (+)110.5 <i>wrong/omitted sign</i> 2 marks (+)184.625 / 184.63 / 184.6 / 185 2 marks <i>No 4CO₂</i> (+)738.5 / 739 <i>No 4CO₂ and no CO/4</i> 1 mark -117.25 / -117.3 / -117 <i>Wrong cycle</i> 2 marks -469 <i>Wrong cycle, no CO/4</i> 1 mark (+)177.875 / 177.88 / 177.9 / 178 1 mark <i>Wrong cycle, no 4CO₂</i> -360.5 <i>Used 118.5</i> 2 marks
		Any other number: CHECK for ECF from 1st marking point for expressions using ALL values with ONE error only e.g. one transcription error:, e.g. 395.3 for 393.5

